Kinetic molecular theory pogil 2005 answers

PHASE CHANGES

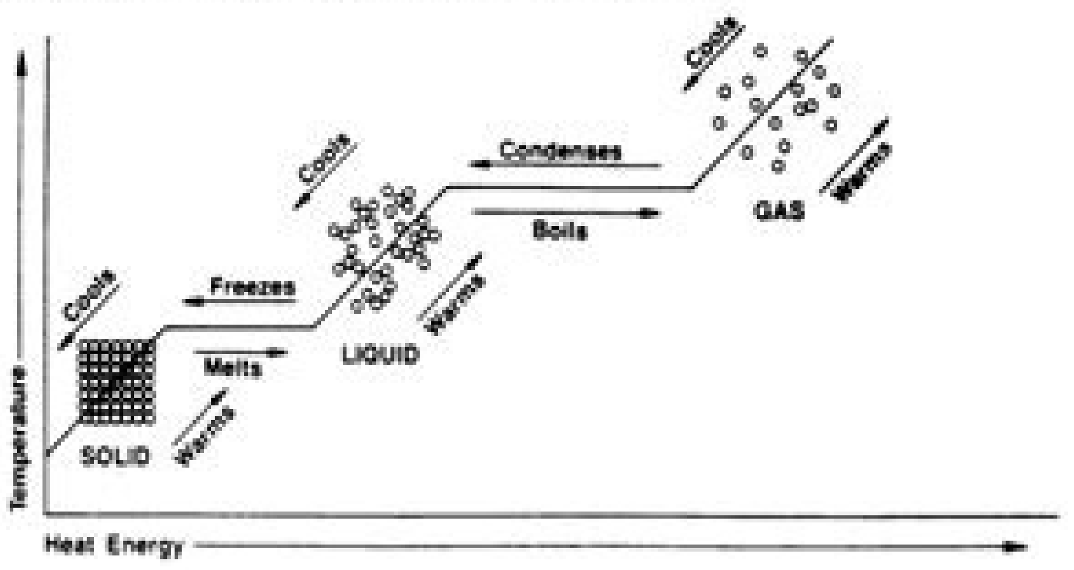
Substances can change from one phase to another. When they do, energy (usually heat) is gained or lost. In this way, solids turn to liquids and liquids to gases when heat energy is gained or absorbed. When heat energy is lost (given off) gases change to liquids and liquids change to solids. All phase changes require a gain or a loss of heat energy. The energy change allows the particles to have a new arrangement, thereby creating a new phase. These phase changes can only happen at certain fixed temperatures for each type of substance.

Changes in phase from solid to liquid (**melting**) and from liquid to gas (boiling) require energy. When solid ice melts and becomes a liquid, the particles of the substance move farther apart and heat energy is gained. When water boils, it forms steam (a gas). The change from liquid water to steam (a gas) is a change in phase and requires the gain of heat energy. This energy can be gained (taken in) from the environment. When you put rubbing alcohol on your skin, it makes your skin feel cold. Your skin feels cold because, when the alcohol changes from a liquid to a gas, it absorbs heat energy from your skin, and the alcohol molecules move further apart. This process is called evaporation. **Evaporation** takes place when liquids turn to gases. Heat energy must be added to the liquid.

Phase changes that require a loss in energy are condensation and freezing. When a liquid becomes a solid (**freeze**), heat energy is generally lost (given off). Energy is also released when a gas becomes a liquid. Condensation happens when gases turn to liquids. Heat energy is given off. The particles slow down, and a liquid forms. Water vapor in the air condenses to form clouds. The droplets of water seen on the outside of a pitcher of cold fruit juice come from water vapor in the air. Water vapor cools enough to condense and collect on the pitcher.

Another type of phase change occurs when a solid changes directly into a gas. This is called **sublimation**. It requires a gain in heat energy. "Dry ice"(solid carbon dioxide) never turns to a liquid before it turns to a gas. Moth balls sublimate making them safe for use next to clothing in storage trunks and closets.

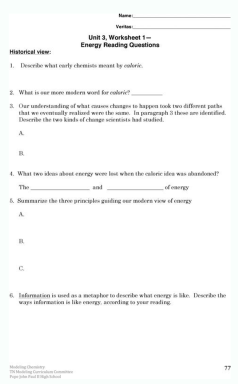
The accompanying graph shows the relationship between temperature and heat energy during the phase changes of water. Study the graph and answer the questions.



1. Does the temperature increase during melting?__________________________
2. Is energy required for each phase change?_____________________________
3. Can both liquid water and steam exist at 100° C ? __________________________
4. What must be changed, temperature or heat energy, during condensation? ______________________
5. How would you describe the change in the arrangement of the particles in the above diagram, as heat energy and temperature increase?

6. What rule can you state about the relationship between phase changes and temperature?

Between phase changes and heat energy?


4. What two ideas about energy were lost when the caloric idea was abandoned?

The ______________ and ______________ of energy

5. Summarize the three principles guiding our modern view of energy

A.

B.

C.

6. Information is used as a metaphor to describe what energy is like. Describe the ways information is like energy, according to your reading.

Pogil 2005 kinetic molecular theory answers. Kinetic molecular theory pogil answer key. What does the kinetic molecular theory explain (at least 3 things). 4 kinetic molecular theory. What are the 5 components of kinetic molecular theory.

Something went wrong. Wait a moment and try again. The requested URL was not found on this server. Additionally, a 404 Not Found error was encountered while trying to use an ErrorDocument to handle the request. Apache/2.4.41 (Ubuntu) Server at senaesteveslab.umassmed.edu Port 443 Standards addressed: California State Standards for Chemistry 3. The conservation of atoms in chemical reactions leads to the principle of conservation of matter and the ability to calculate the mass of products and reactants. As a basis for understanding this concept: e. Students know how to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses. 4. The kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases. As a basis for understanding this concept: b Students know the random motion of molecules explains the diffusion of gases. 6. Solutions are homogenous mixtures of two or more substances. As a basis for understanding this concept: b. Students know how to describe the dissolving process at the molecular level by using the concept of random molecular motion State the postulates of the kinetic-molecular theory Use this theory's postulates to explain the gas laws The gas laws that we have seen to this point, as well as the ideal gas equation, are empirical, that is, they have been derived from experimental observations. The mathematical forms of these laws closely describe the macroscopic behavior of most gases at pressures less than about 1 or 2 atm. Although the gas laws describe relationships that have been verified by many experiments, they do not tell us why gases follow these relationships. The kinetic molecular theory (KMT) is a simple microscopic model that effectively explains the gas laws described in previous modules of this chapter. This theory is based on the following five postulates described here. (Note: The term "molecule" will be used to refer to the individual chemical species that compose the gas, although some gases are composed of atomic species, for example, the noble gases.) Gases are composed of molecules that are in continuous motion, travelling in straight lines and changing direction only when they collide with other molecules or with the walls of a container. The molecules composing the gas are negligibly small compared to the distances between them. The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls. Gas molecules exert no attractive or repulsive forces on each other or the container walls; therefore, their collisions are elastic (do not involve a loss of energy). The average kinetic energy of the gas molecules is proportional to the kelvin temperature of the gas. The test of the KMT and its postulates is its ability to explain and describe the behavior of a gas. The various gas laws can be derived from the assumptions of the KMT, which have led chemists to believe that the assumptions of the theory accurately represent the properties of gas molecules. We will first look at the individual gas laws (Boyle's, Charles's, Amontons's, Avogadro's, and Dalton's laws) conceptually to see how the KMT explains them. Then, we will more carefully consider the relationships between molecular masses, speeds, and kinetic energies with temperature, and explain Graham's law. Recalling that gas pressure is exerted by rapidly moving gas molecules and depends directly on the number of molecules hitting a unit area of the wall per unit of time, we see that the KMT conceptually explains the behavior of a gas as follows: Amontons's law. If the temperature is increased, the average speed and kinetic energy of the gas molecules increase. If the volume is held constant, the increased speed of the gas molecules results in more frequent and more forceful collisions with the walls of the container, therefore increasing the pressure (Figure 9.31). Charles's law. If the temperature of a gas is increased, a constant pressure may be maintained only if the volume occupied by the gas increases. This will result in greater average distances traveled by the molecules to reach the container walls, as well as increased wall surface area. These conditions will decrease the both the frequency of molecule-wall collisions and the number of collisions per unit area, the combined effects of which balance the effect of increased collision forces due to the greater kinetic energy at the higher temperature. Boyle's law. If the gas volume volume of a given amount of gas at a given temperature is decreased (that is, if the gas is compressed), the molecules will be exposed to a decreased container wall area. Collisions with the container wall will therefore occur more frequently and the pressure exerted by the gas will increase (Figure 9.31). Avogadro's law. At constant pressure and temperature, the frequency and force of molecule-wall collisions are constant. Under such conditions, increasing the number of gaseous molecules will require a proportional increase in the container volume in order to yield a decrease in the number of collisions per unit area to compensate for the increased frequency of collisions (Figure 9.31). Dalton's Law. Because of the large distances between them, the molecules of one gas in a mixture bombard the container walls with the same frequency whether other gases are present or not, and the total pressure of a gas mixture equals the sum of the (partial) pressures of the individual gases. Figure 9.31 (a) When gas temperature increases, gas pressure increases due to increased force and frequency of molecular collisions. (b) When volume decreases, gas pressure increases due to increased frequency of molecular collisions. (c) When the amount of gas increases at a constant pressure, volume increases to yield a constant number of collisions per unit wall area per unit time. The previous discussion showed that the KMT qualitatively explains the behaviors described by the various gas laws. The postulates of this theory may be applied in a more quantitative fashion to derive these individual laws. To do this, we must first look at speeds and kinetic energies of gas molecules, and the temperature of a gas sample. In a gas sample, individual molecules have widely varying speeds; however, because of the vast number of molecules and collisions involved, the molecular speed distribution and average speed are constant. This molecular speed distribution is known as a Maxwell-Boltzmann distribution, and it depicts the relative numbers of molecules in a bulk sample of gas that possesses a given speed (Figure 9.32). Figure 9.32 The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, νp, is a little less than 400 m/s, while the root mean square speed, urms, is closer to 500 m/s. The kinetic energy (KE) of a particle of mass (m) and speed (u) is given by: KE=12mu2KE=12mu2 Expressing mass in kilograms and speed in meters per second will yield energy values in units of joules (J = kg m2 s–2). To deal with a large number of gas molecules, we use averages for both speed and kinetic energy. In the KMT, the root mean square speed of a particle, urms, is defined as the square root of the average of the squares of the speeds with n = the number of particles: urms=u2¯=u12+u22+u32+u42+…nurms=u2¯=u12+u22+u32+u42+…nThe average kinetic energy for a mole of particles, KEavg, is then equal to:KEavg=12Murms2KEavg=12Murms2where M is the molar mass expressed in units of kg/mol. The KEavg of a mole of gas molecules is also directly proportional to the temperature of the gas and may be described by the equation:KEavg=32RTKEavg=32RT where R is the gas constant and T is the kelvin temperature. When used in this equation, the appropriate form of the gas constant is 8.314 J/mol·K (8.314 kg m2s–2mol–1K–1). These two separate equations for KEavg may be combined and rearranged to yield a relation between molecular speed and temperature:12Murms2=32RT12Murms2=32RTurms=3RTMurms=3RTM Calculation of urmsCalculate the root-mean-square speed for a nitrogen molecule at 30 °C.Solution Convert the temperature into Kelvin: 30°C+273=303 K30°C+273=303 K Determine the molar mass of nitrogen in kilograms:28.0g1 mol×1 kg1000g=0.028kg/mol28.0g1 mol×1 kg1000g=0.028kg/mol Replace the variables and constants in the root-mean-square speed equation, replacing Joules with the equivalent kg m2s–2: urms=3(8.314J/mol K)(303 K)(0.028kg/mol)=2.70×105m2s−2=519m/surms=3(8.314J/mol K)(303 K)(0.028kg/mol)=2.70×105m2s−2=519m/sCheck Your LearningCalculate the root-mean-square speed for a mole of oxygen molecules at –23 °C. If the temperature of a gas increases, its KEavg increases, more molecules have higher speeds and fewer molecules have lower speeds, and the distribution shifts toward higher speeds overall, that is, to the right. If temperature decreases, KEavg decreases, more molecules have lower speeds and fewer molecules have higher speeds, and the distribution shifts toward lower speeds overall, that is, to the left. This behavior is illustrated for nitrogen gas in Figure 9.33. Figure 9.33 The molecular speed distribution for nitrogen gas (N2) shifts to the right and flattens as the temperature increases; it shifts to the left and heightens as the temperature decreases. At a given temperature, all gases have the same KEavg for their molecules. Gases composed of lighter molecules have more high-speed particles and a higher urms, with a speed distribution that peaks at relatively higher speeds. Gases consisting of heavier molecules have more low-speed particles, a lower urms, and a speed distribution that peaks at relatively lower speeds. This trend is demonstrated by the data for a series of noble gases shown in Figure 9.34. Figure 9.34 molecular speed is directly related to molecular mass. At a given temperature, lighter molecules move faster on average than heavier molecules. The gas simulator may be used to examine the effect of temperature on molecular speeds. Examine the simulator's "energy histograms" (molecular speed distributions) and "species information" (which gives average speed values) for molecules of different masses at various temperatures. According to Graham's law, the molecules of a gas are in rapid motion and the molecules themselves are small. The average distance between the molecules of a gas is large compared to the size of the molecules. As a consequence, gas molecules can move past each other easily and diffuse at relatively fast rates. The rate of effusion of a gas depends directly on the (average) speed of its molecules: effusion rate∝urmseffusion rate∝urms Using this relation, and the equation relating molecular speed to mass, Graham's law may be easily derived as shown here: urms=3RTMurms=3RTMM=3RTurms2=3RTu¯2M=3RTurms2=3RTu¯2effusion rate Aeffusion rate B=urmsAurmsB=3RTMA3RTMB=MBMAeffusion rate Aeffusion rate B=urmsAurmsB=3RTMA3RTMB=MBMAThe ratio of the rates of effusion is thus derived to be inversely proportional to the ratio of the square roots of their masses. This is the same relation observed experimentally and expressed as Graham's law.

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